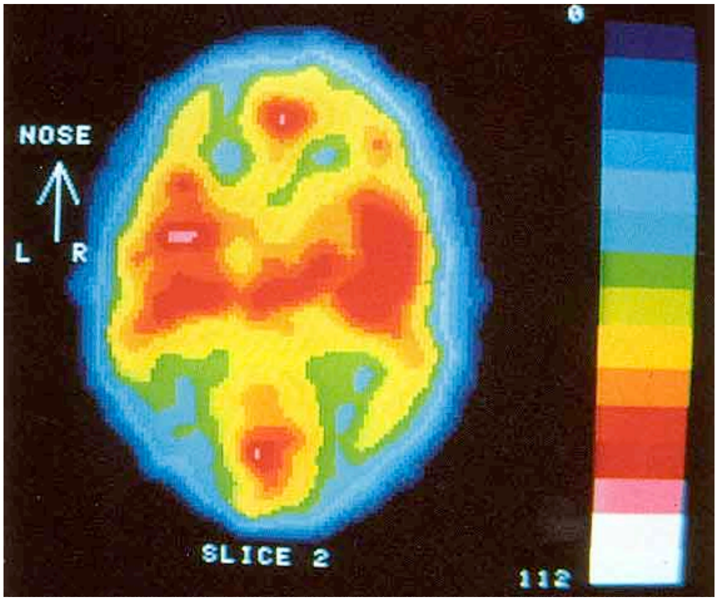
Chapter 2: Elements and Compounds



Chapter 2
2.1 The Structure of the Atom
2.2 Elements and the Periodic Table
2.3 Covalent Compounds
2.4 Ions and Ionic Compounds

Chapter Goals
- Recognize the components of an atom.
- Describe the organization of the periodic table.
- Identify covalent compounds and their properties.
- Understand the relationship between atoms and ions.
- Identify ionic compounds and their properties.

Chapter In Context

As we learned in Chapter 1, each of the 116 known elements is composed of a unique type of atom. In this chapter, we discover that there are actually a number of different variations, or isotopes, of the atom associated with each element. We explore the structure of the atom in further detail and learn how different isotopes are formed. We also discuss the different ways that molecules and compounds are represented and named.

Atoms make up all the matter around us, so an understanding of atoms, and particularly how we can manipulate atoms to affect their properties, has led to numerous benefits in science and in our everyday lives.

- Biology: Over the last sixty years, since we began to synthesize new elements in the laboratory, synthetic isotopes have found significant use in medicine. For example, when compounds containing the synthetic isotope technetium, are injected into patients, portions of the body such as the brain, the heart, and the circulatory system can be imaged.

2.1 The Structure of the Atom

OWL Opening Exploration 2.1

Atoms, the smallest unit of matter, consist of three subatomic particles: protons, neutrons, and electrons. A **proton** carries a relative charge of +1 and has a mass of 1.672622×10^{-24} g. A **neutron** carries no electrical charge and has a mass of 1.674927×10^{-24} g. An **electron** carries a relative charge of –1 and has a mass of 9.109383×10^{-27} g. Two of the subatomic particles, protons and neutrons, are found in the **atomic nucleus**, a very small region of high density at the center of the atom. Electrons are found in the region around the nucleus. As you will see when we study atomic structure in more detail in an upcoming chapter, the precise location of electrons is not determined. Instead, we visualize an electron cloud surrounding the nucleus that represents the most probable location of electrons (Figure 2.1). Note that the atom represented in Figure 2.1 is not drawn to scale. In reality, electrons account for most of the volume of an atom, and the nucleus of an atom is about one ten-thousandth (1/10,000) the diameter of a typical atom. For example, if you expand an atom so that it has a diameter the same size as a football field, 100 yards or about 90 meters, the nucleus of the atom would have a diameter of only about 1 cm!

Flashforward
Chapter 7, Atomic Structure

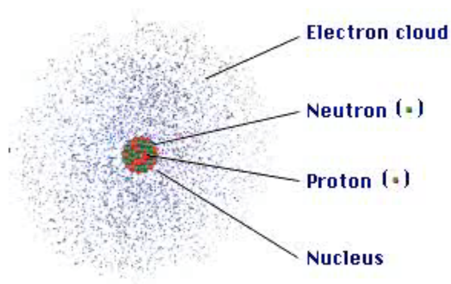


Figure 2.1 The arrangement of subatomic particles in an atom (not drawn to scale).

The mass and charge of an atom affects the physical and chemical properties of the element and the compounds it forms. As shown in Table 2.1, the three subatomic particles are easily differentiated by both charge and mass.

Table 2.1 Properties of Subatomic Particles

	Actual mass (kg)	Relative mass	Mass (u)	Actual charge (C)	Relative charge
Proton (p)	1.672622×10^{-27}	1836	1.007276	1.602×10^{-19} C	+1
Neutron (n)	1.674927×10^{-27}	1839	1.008665	0	0
Electron (e^-)	9.109383×10^{-31}	1	5.485799×10^{-4} (≈ 0)	-1.602×10^{-19} C	–1

The mass of an atom is almost entirely accounted for by its dense nucleus of protons and neutrons. The actual mass of protons and neutrons is very small, so it is more convenient to define the mass of these particles using a different unit. The **atomic mass unit** (**u**) is defined as one-twelfth of the mass of a carbon atom that contains 6 protons and 6 neutrons. Because neutrons and protons have very similar masses, both have a mass of approximately 1 u. The mass of an electron is about 2000 times less that of protons and neutrons, and it has a mass of approximately zero on the atomic mass unit scale. Putting together all of this information, we can now represent a typical atom more accurately, as shown in Figure 2.2.

Figure to come

Figure 2.2 Diagram of a carbon atom.

Atoms are neutral because protons and electrons have equal, opposite charges and atoms have equal numbers of positively charge protons and negatively charged electrons. An **ion** is an atom with an unequal number of protons and electrons, which therefore carries an overall positive or negative charge. When an atom carries more protons than electrons, it carries an overall positive charge, and is called a **cation**. An atom with more electrons than protons has an overall negative charge and is called an **anion**. As you will see later in this chapter, ions have very different properties than the elements they are derived from.

Chapter Goals Revisited
- Recognize the components of an atom. **Write atomic symbols.**

Atomic Number, Mass Number, and Atomic Symbols

Atoms of each element can be distinguished by the number of protons in the nucleus. The **atomic number** (***Z***) of an element is equal to the number of protons in the nucleus. For example, the carbon atom shown in Figure 2.2 has six protons in its nucleus, and therefore carbon has an atomic number of six ($Z = 6$). Each element has a unique atomic number and all atoms of that element have the same number of protons in the nucleus. All atoms of hydrogen have one proton in the nucleus ($Z = 1$) and all atoms of gold have 79 protons in the nucleus ($Z = 79$). Because protons carry a positive charge (+1), in a neutral atom the atomic number also equals the number of electrons (–1) in that atom.

An atom can also be characterized by its mass. Because the mass of electrons is negligible, the mass of an atom in atomic mass units (u) is essentially equal to the number of protons and neutrons in the nucleus of an atom, called the **mass number** (***A***). For example, a carbon atom with 6 protons and 6 neutrons in its nucleus has a mass number of 12 ($A = 12$), and a gold atom with 79 protons and 119 neutrons in its nucleus has a mass number of 198 ($A = 198$).

The **atomic symbol** for an element consists of the one- or two-letter symbol that represents the element along with the atomic number in subscript and the mass number in superscript (Figure 2.3). For example, the atomic symbol for a carbon (C) atom with 6 protons and 6 neutrons is ${}^{12}_{6}\mathrm{C}$, and the symbol for a gold (Au) atom with 79 protons and 119 neutrons is ${}^{198}_{79}\mathrm{Au}$. Note that the number of neutrons in an atom is equal to the difference between the mass number and the atomic number.

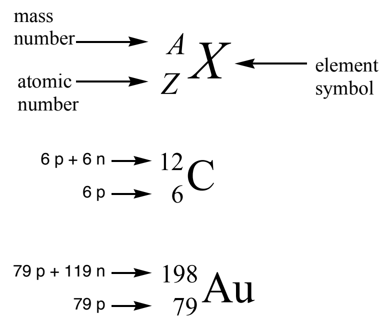


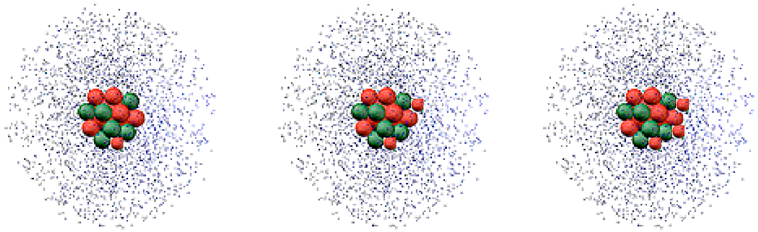
Figure 2.3 Atomic symbols

EXAMPLE PROBLEM: Atomic Symbols
Write the atomic symbol for the following atoms.
(a) A nitrogen atom containing 7 protons, 8 neutrons, and 7 electrons.
(b) A uranium atom containing 92 protons, 143 neutrons, and 92 electrons.
SOLUTION:
(a) ${}^{15}_{7}\mathrm{N}$ The atomic number of nitrogen is equal to the number of protons (7) and the mass number is equal to the number of protons plus the number of neutrons ($7 + 8 = 15$).
(b) ${}^{235}_{92}\mathrm{U}$ The atomic number of uranium is equal to the number of protons (92) and the mass number is equal to the number of protons plus the number of neutrons ($92 + 143 = 235$).

OWL Example Problems
2.2 Atomic Symbols: Tutor
2.3 Atomic symbols

Isotopes and Atomic Mass

Whereas all atoms of a given element have the same number of protons, the number of neutrons found in the nucleus for a particular element can vary. For example, while atoms of carbon always have 6 protons in the nucleus, a carbon atom might have 6, 7 or even 8 neutrons in the nucleus (Figure 2.3). Even though the mass number for these atoms differs, each is a carbon atom because each has 6 protons in the nucleus.



6 protons 6 neutrons

6 protons 7 neutrons

6 protons 8 neutrons

Figure 2.3 Isotopes of carbon

Every carbon atom has 6 protons and the mass of electrons is negligible, so we can conclude that the carbon atoms shown in Figure 2.4 have different mass numbers because each has a different number of neutrons. Atoms that have the same atomic number (Z) but different mass numbers (A) are called **isotopes**. Isotopes are named using the element name and the mass number. For example, the isotopes shown in Figure 2.4 are named carbon-12, carbon-13, and carbon-14. The atomic symbols for these elements can be written ^{12}C, ^{13}C, and ^{14}C. Notice that because the atomic number is always the same for a given element, Z is sometimes omitted from the atomic symbol for an isotope.

Whereas some elements such as fluorine have only one naturally occurring isotope, other elements have as many as ten. Most samples of an element in nature contain a mixture of various isotopes. When we talk about the mass of an atom of a certain element, therefore, we must take into account that any sample of that element would include different isotopes with different masses. The **atomic mass** (sometimes called *atomic weight*) for any element is the average mass of all naturally occurring isotopes of that element, taking into account the relative abundance of the isotopes. For example, chlorine (Z = 17) has two naturally occurring isotopes, ^{35}Cl and ^{37}Cl. In any sample of chlorine, about ¾ of the atoms are ^{35}Cl and about ¼ are ^{37}Cl. Because there are more ^{35}Cl atoms than ^{37}Cl atoms in the sample, the average of mass of chlorine is closer to that of ^{35}Cl than to that of ^{37}Cl. Atomic mass is therefore a weighted average of the atomic masses of all isotopes for a particular element.

Chapter Goals Revisited

- Recognize the components of an atom. **Use atomic mass to identify isotopes of an element.**

Average atomic mass depends on both the mass of each isotope present and the relative abundance of that isotope. In order to calculate the average atomic mass for an element, the fractional abundance and the exact mass of the isotopes are summed as shown in Equation 2.1 and in the following example.

$$\text{Average atomic mass} = \sum_{\substack{\text{all}\\ \text{isotopes}}} (\text{exact mass})(\text{fractional abundance}) \qquad \textbf{2.1}$$

EXAMPLE PROBLEM: Average Atomic Mass

Calculate the average atomic mass for chlorine. Chlorine has two naturally occurring isotopes, chlorine-35 (34.96885 u, 75.78% abundant) and chlorine-37 (36.96590 u, 24.22% abundant).

SOLUTION:

Use Equation 2.1, the exact mass of the isotopes, and the fractional abundance of the isotopes to calculate the average atomic mass of chlorine.

Average atomic mass (Cl) = (^{35}Cl exact mass)(^{35}Cl fractional abundance) + (^{37}Cl exact mass)(^{37}Cl fractional abundance)

$$\text{Average atomic mass (Cl)} = (34.96885\text{ u})\left(\frac{75.78}{100}\right) + (36.96560\text{ u})\left(\frac{24.22}{100}\right)$$

Average atomic mass (Cl) = 35.45 u

Note that the average atomic mass of chlorine is closer to the mass of ^{35}Cl, the more abundant isotope, than to that of ^{37}Cl.

> *OWL Example Problems*
> *2.4 Average Atomic Mass: Tutor*
> *2.5 Average Atomic Mass*

2.2 Elements and the Periodic Table

> *OWL Opening Exploration*
> *2.6*

Chapter Goals Revisited
- Describe the organization of the periodic table.

Information about the elements is organized in the **periodic table of the elements** (Figure 2.4 and inside the front cover of this textbook). The periodic table is the most important tool that chemists use. It not only contains information specific to each element but it also organizes the elements according to their physical and chemical properties.

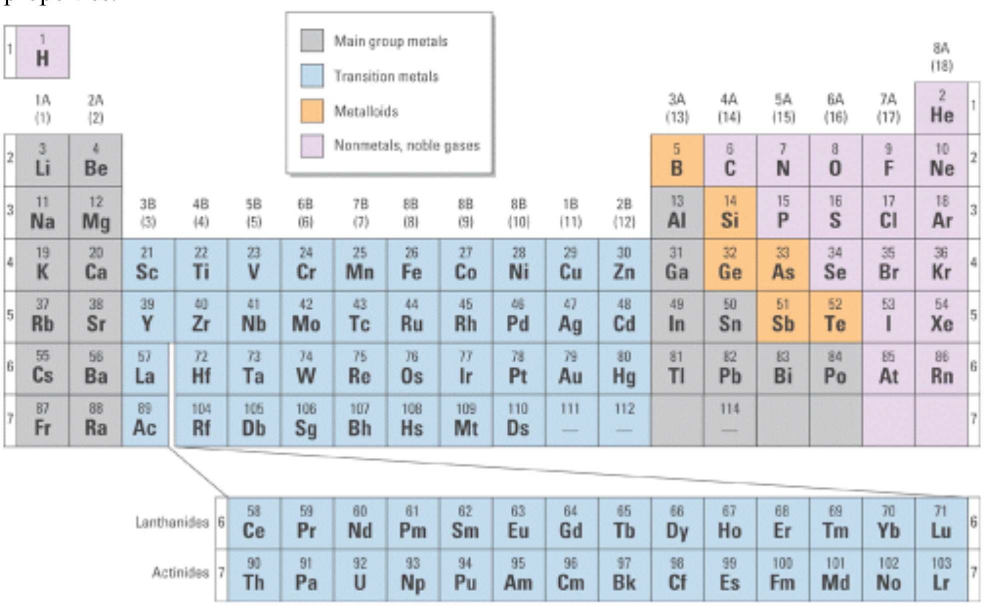


Figure 2.4 The periodic table

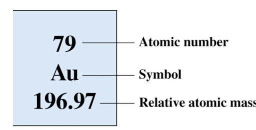


Figure 2.5 The periodic table entry for silver

Each entry in the periodic table represents a single element (Figure 2.5) and contains the element's chemical symbol, atomic number and average atomic mass. The elements are arranged in vertical columns called **groups** and horizontal rows called **periods**. The elements within each group have similar chemical and physical properties. The periodic, repeating properties of the elements within groups is one of the most important aspects of the periodic table, as you will see in an upcoming chapter. Some of the groups in the periodic table are given special names (Table 2.2) to reflect their common properties. For example, the Group 1A elements, the alkali metals, are all shiny solids that react vigorously with air, water and the Group 7A elements, the halogens.

Flash-forward
Chapter X
Periodic trends

Table 2.2 The special names for some groups in the periodic table

Group	Name
1A	Alkali metals
2A	Alkaline earth metals
7A	Halogens
8A	Noble gases

The eighteen groups in the periodic table are numbered according to one of three common numbering schemes. The scheme shown in Figure 2.x is widely used in North America, and consists of a number followed by A or B. The elements in A groups are the **main group** elements, also called the *representative* elements, and the elements in B groups are **transition elements**. The International Union of Pure and Applied Chemistry (IUPAC) has proposed a simpler numbering scheme, also shown in Figure 2.x, that numbers the groups 1–18 from left to right.

There are 7 horizontal periods in the periodic table. Portions of periods 6 and 7 are placed below the main body of the periodic table in order to make it fit easily on a page. These portions of periods 6 and 7 are given special names, the **lanthanides** and **actinides**.

The elements on the left side of the periodic table are **metals**, the elements on the right side are **nonmetals**, and the elements at the interface of these two regions are **metalloids** or **semi-metals**. Metals are generally shiny solids (mercury, Hg, is the only liquid metal at room temperature) that are ductile and good conductors of electricity. Nonmetals are generally dull, brittle solids or gases (bromine is the only liquid nonmetal at room temperature) that do not conduct electricity. Metalloids have properties of both metals and nonmetals.

Most of the elements in the periodic table are solids. Only two elements exist as liquids at room temperature (mercury and bromine) and eleven elements are gases at room temperature (hydrogen, nitrogen, oxygen, fluorine, chlorine, helium, neon, argon, krypton, xenon, and radon). At the atomic level, elements are found as individual atoms (helium, He, sodium, Na, and mercury, Hg, for example), as molecules made up of two or more atoms of an element (oxygen, O_2, sulfur, S_8, and white phosphorus, P_4, for example), or as a connected three-dimensional array of atoms (silicon, Si, carbon, C, red phosphorus, P).

Red phosphorus, P, and white phosphorus, P_4, are examples of **allotropes**, forms of the same element that differ in their physical and chemical properties. Red phosphorus, which consists of long chains of phosphorus atoms, is nontoxic, has a deep red color, and burns in air at high temperatures (above 250 ºC). White phosphorus, which is made up of individual molecules of four phosphorus atoms, is a white or yellow waxy solid that ignites in air above 50 ºC and is very poisonous. Other examples of elements that exist as different allotropes are oxygen (diatomic oxygen, O_2, and triatomic ozone, O_3) and carbon (diamond, graphite, and buckminsterfullerene).

The periodic table is used extensively in chemistry, and it is helpful to become familiar with the structure of the table. You should learn the names and symbols for the first 36 elements and some other common elements such as silver (Ag), gold (Au), tin (Sn), iodine (I), lead (Pb), and uranium (U).

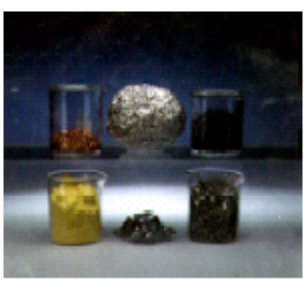


Figure 2.6 Some common elements

OWL Example Problems
2.7 The Periodic Table: Exercise
2.8a The Periodic Table
2.8b Element symbols and the Periodic Table
2. 9 Molecules and Elements

2.3 Covalent Compounds

OWL Opening Exploration 2.10

> **Chapter Goals Revisited**
> - Identify covalent compounds and their properties.
>
>   **Use formulas and models to represent covalent compounds.**

As we learned in Chapter 1, compounds are formed when two or more elements are combined chemically in a defined ratio. **Covalent compounds** consist of atoms of different elements held together by covalent bonds. We will discuss the details of covalent bonds in Chapter 8. Water, H_2O, is an example of a **molecular covalent compound**. Water is made up of individual H_2O molecules, with the oxygen and hydrogen atoms in each water molecule held together by covalent bonds (Figure 2.7a).

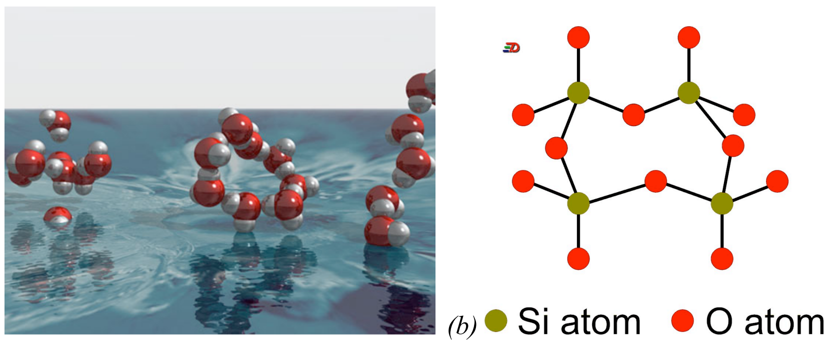
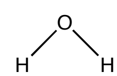
(a) *(b)*


Figure 2.7 (a) Water molecules and (b) Si-O bonded network in sand

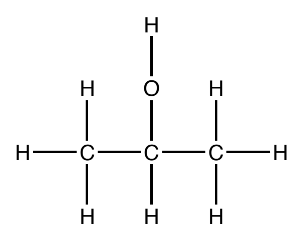
Silicon dioxide, SiO_2, also known as sand, is an example of a **network covalent compound**. Unlike water, which consists of individual H_2O molecules, silicon dioxide is made up of a three-dimensional network of silicon and oxygen atoms held together by covalent bonds. (Figure 2.7b).

Representing Covalent Compounds: Molecular and Empirical Formulas

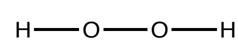
Molecules can vary in complexity from only two atoms to many. The simplest way to represent a molecule is through a **molecular formula**. A molecular formula contains the symbol for each element present and a subscript number to identify the number of atoms of each element the molecule. If only one atom of an element is present in a molecule, however, the number 1 is not used. A water molecule, H_2O, is made up of two hydrogen atoms and one oxygen atom. Isopropanol has the molecular formula C_3H_8O, which means that a single molecule of isopropanol contains three carbon atoms, eight hydrogen atoms, and one oxygen atom. Notice that chemical formulas always show a whole-number ratio of elements.



Water, H_2O



Isopropanol, C_3H_8O



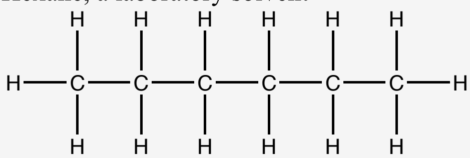
Hydrogen peroxide, H_2O_2

Whereas the molecular formula indicates the number of atoms of each element in one molecule of a compound, an **empirical formula** represents the simplest whole-number ratio of elements in a compound. Hydrogen peroxide, for example, is a molecular compound with the molecular formula of H_2O_2. The empirical formula of hydrogen peroxide, HO, shows the simplest whole-number ratio of elements in the compound. Network covalent compounds are also represented using empirical formulas. Silicon dioxide, for example, does not consist of individual SiO_2 molecules. The simplest ratio of elements in the compound is 1 Si : 2 O atoms. The empirical formula of silicon dioxide is therefore SiO_2. Carbon (diamond) is an example of a network element. It is made up of carbon atoms held together by covalent bonds in a three-dimensional network and the element is represented by the empirical formula C.

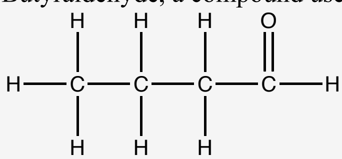
EXAMPLE PROBLEM: Molecular and Empirical Formulas

Determine the molecular and empirical formulas for the substances below.

(a) Hexane, a laboratory solvent



(b) Butyraldehyde, a compound used for synthetic almond flavoring in food



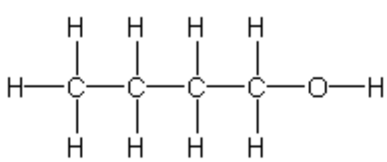
SOLUTION:

(a) A hexane molecule contains 6 carbon atoms and 14 hydrogen atoms. The molecular formula of hexane is C_6H_{14}. The empirical formula, the simplest whole-number ratio of elements in the compound, is C_3H_7.

(b) A butyraldehyde molecule contains 4 carbon atoms, 8 hydrogen atoms and 1 oxygen atom. The molecular formula of butyraldehyde is C_4H_8O. The empirical formula, the simplest whole-number ratio of elements in the compound, is the same as the molecular formula, C_4H_8O.

OWL Example Problems
2.11a Molecular and Empirical Formulas

A molecular formula identifies the number and types of elements present in a molecule, but it does not provide information on how the atoms are connected. A **structural formula** shows the linkage of all the atoms in the molecule. The covalent bonds are represented by lines between the element symbols. Butanol, a molecular compound made up of four carbon atoms, ten hydrogen atoms, and one oxygen atom has the following structural formula:



A **condensed structural formula** lists the atoms present in groups to indicate connectivity between the atoms. The condensed structural formula for butanol is $CH_3CH_2CH_2CH_2OH$. Interpretation of this type of formula requires familiarity with commonly encountered groups of atoms, such as the CH_3 or CH_2 groups. Note that while the structural formula does convey information about connectivity, it does not convey information about the 3-dimensional shape of the compound.

Representing Covalent Compounds: Molecular Models

Chemists often need to visualize the three-dimensional shape of a molecule in order to understand its chemical or physical properties. A variety of models are used to represent the shapes of molecules, each of which has a different purpose. The **wedge and dash model**, shown in Figure 2.8(a) is a two-dimensional representation of a three-dimensional structure that can easily be drawn on paper. In this model of molecular shape, bonds are represented by lines (bonds that lie in the plane of the paper), wedges (bonds that lie in front of the plane of the paper), or dashes (bonds that lie behind the plane of the paper).

The other two common types of molecular models are created using molecular modeling software, sophisticated computer programs that calculate the spatial arrangement of atoms and bonds. A **ball and stick model** shows atoms as colored spheres connected by sticks that represent covalent bonds (Figure 2.8(b)). This type of model emphasizes the connections between atoms and the arrangement of atoms in the molecule. A less accurate ball and stick model can be created using a commercial molecular modeling kit. In this case, the model can be held in your hands and the atoms and bonds can be rotated.

In a **space-filling model**, shown in Figure 2.8(c), interpenetrating spheres represent the relative amount of space occupied by each atom in the molecule. This type of model is useful when considering the overall shape of molecules and how molecules interact when they come in contact with one another.

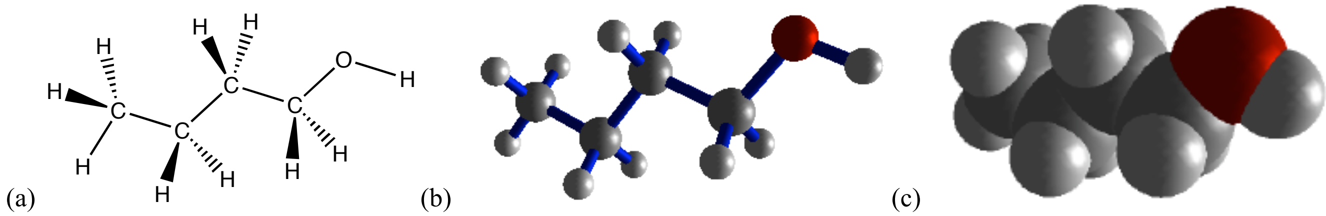


Figure 2.8 Wedge and dash (a), ball and stick (b), and space-filling (c) models of butanol

OWL Example Problems
2.11b Representing Compounds: Tutor
2.11c Representing Molecular Compounds

Naming Covalent Compounds

There are many ways to categorize covalent compounds, and two common classes are **binary nonmetals** and **inorganic acids**. Often, a compound will belong to more than one of these categories. Covalent compounds are named according to guidelines created by the Chemical Nomenclature and Structure Representation Division of IUPAC, the International Union of Pure and Applied Chemistry. Some compounds, however, have names that do not follow these guidelines because they have been known by other common names for many years.

Naming Covalent Compounds: Binary Nonmetals

Binary nonmetal compounds consist of only two elements, both nonmetals. Some examples include H_2O, CS_2, and SiO_2.

Binary nonmetal compounds are named according to the following steps:

1. The first word in the compound name is the name of the first element in the compound formula. If the compound contains more than one atom of the first element, use a prefix (Table 2.3) to indicate the number of atoms in the formula.
   Example: CS_2 First word in compound name: carbon
   N_2O_4 First word in compound name: dinitrogen
2. The second word in the compound name is the name of the second element in the formula that has been changed to end with *–ide*. In all cases, use a prefix (Table 2.x) to indicate the number of atoms in the formula.
   Example: CS_2 Second word in compound name: disulfide
   N_2O_4 Second word in compound name: tetraoxide
3. Name the compound by combining the first and second words of the compound name.
   Example: CS_2 carbon disulfide
   N_2O_4 dinitrogen tetraoxide

Chapter Goals Revisited

- Identify covalent compounds and their properties.
  Name covalent compounds.

Table 2.3 Prefixes used in naming binary nonmetal compounds

Number	Prefix
1	mono–
2	di–
3	tri–
4	tetra–
5	penta–
6	hexa–
7	hepta
8	octa–
9	nona–
10	deca–
12	dodeca–

Many binary nonmetal compounds have special names that have been used for many years. Examples include water (H_2O), ammonia (NH_3), and nitric oxide (NO). **Hydrocarbons**, binary nonmetal compounds containing only carbon and hydrogen, are also given special names. These compounds are one class of organic compounds, compounds that contain carbon and hydrogen and often other elements such as oxygen and nitrogen. Hydrocarbons are named according to the number of carbon and hydrogen atoms in the compound formula, as shown in Table 2.4. The names of some common binary nonmetal compounds are shown in Table 2.5.

Table 2.4 Hydrocarbons with the formula C_nH_{2n+2}

Hydrocarbon	Name
CH_4	methane
C_2H_6	ethane
C_3H_8	propane
C_4H_{10}	butane
C_5H_{12}	pentane
C_6H_{14}	hexane
C_8H_{18}	octane

Table 2.5 Names and formulas of some binary nonmetals

Name	Formula	Name	Formula
water	H_2O	sulfur dioxide	SO_2
hydrogen peroxide	H_2O_2	sulfur trioxide	SO_3
ammonia	NH_3	carbon monoxide	CO
hydrazine	N_2H_4	carbon dioxide	CO_2
nitric oxide	NO	chlorine monoxide	ClO
nitrogen dioxide	NO_2	disulfur decafluoride	S_2F_{10}

Naming Covalent Compounds: Inorganic Acids

Inorganic acids produce the hydrogen ion (H^+) when dissolved in water, and are compounds that contain hydrogen and one or more nonmetals. Inorganic acids can often be identified by their chemical formulas because hydrogen is the first element in the compound formula. Some examples include HCl, H_2S, and HNO_3.

Inorganic acids are named as binary nonmetal compounds but without the use of prefixes (HCl, hydrogen chloride, H_2S, hydrogen sulfide), or using common names (HNO_3, nitric acid, H_2SO_4, sulfuric acid). Groups of acids that differ only in the number of oxygen atoms, **oxoacids**, are named according to the number of oxygen atoms in the formula. Chlorine, bromine, and iodine each form a series of four oxoacids, as shown in Table 2.6.

Table 2.6 Names and formulas of the halogen Oxoacids

Formula	Name	Formula	Name	Formula	Name
$HClO_4$	perchloric acid	$HBrO_4$	perbromic acid	HIO_4	periodic acid
$HClO_3$	chloric acid	$HBrO_3$	bromic acid	HIO_3	iodic acid
$HClO_2$	chlorous acid	$HBrO_2$	bromous acid	HIO_2	iodous acid
HClO	hypochlorous acid	HBrO	hypobromous acid	HIO	hypoiodous acid

When naming oxoacids, the suffix *–ic* is generally used to indicate an acid with more oxygen atoms and the suffix *–ous* is used to indicate an acid with fewer oxygens. For example, HNO_3 is nitric acid and HNO_2 is nitrous acid, H_2SO_4 is sulfuric acid and H_2SO_3 is sulfurous acid. Some common acids are shown in Table 2.7.

Table 2.7 Names and formulas of some inorganic acids

Name	Formula	Name	Formula
hydrogen chloride	HCl	nitric acid	HNO_3
hydrogen bromide	HBr	nitrous acid	HNO_2
hydrogen sulfide	H_2S	sulfuric acid	H_2SO_4
phosphoric acid	H_3PO_4	sulfurous acid	H_2SO_3

EXAMPLE PROBLEM: Naming Covalent Compounds

Name or write the formula for the following covalent compounds:

(a) CF_4 (b) P_4S_3 (c) hydrogen iodide (d) hydrazine

SOLUTION:

(a) Carbon tetrafluoride — Notice that the name of the first element, carbon, does not include the *mono–* prefix.
(b) Tetraphosphorus trisulfide — Both element names include prefixes and the name of the second element ends in *–ide*.
(c) HI — This is the formula of an inorganic acid.
(d) N_2H_4 — This is a common name that must be memorized.

OWL Example Problems
2.12 Naming Binary Covalent Compounds: Tutor
2.13a Naming Binary Covalent Compounds
2.13b Naming Inorganic Acids

2.4 Ions and Ionic Compounds

OWL Opening Exploration
2.14 Opening Exploration: Ionic Compounds

Unlike covalent compounds, **ionic compounds** contain ions (Section 2.1), species that carry a positive (cation) or negative (anion) charge. While the atoms in covalent compounds are held together by covalent bonds, ionic compounds are held together by strong attractive forces between cations and anions. The different makeup of these two types of compounds results in species with very different physical and chemical properties. For example, many covalent compounds are gases, liquids, or solids with low melting points while most ionic compounds are solids with very high melting points.

Chapter Goals Revisited
- Understand the relationship between atoms and ions. **Predict charges on monoatomic main group elements.**

Monoatomic Ions

Remember that an atom carries no charge because it contains an equal number of positively charged protons and negatively charged electrons. When a single atom gains or loses one or more electrons, the number of electrons and protons is no longer equal and a **monoatomic ion** is formed. The charge on an ion is indicated using a superscript to the right of the element symbol. When the charge is +1 or –1, it is written without the number 1. For example, magnesium forms a cation, Mg^{2+}, when it loses two electrons, and bromine forms an anion, Br^{-}, when it gains one electron (Figure 2.9).

Ion Charges
Notice that when an ion charge is written with an atom symbol, the numeric value is written first, followed by the + or – symbol. When describing the charge on an ion, the + or – symbol is written first, followed by the numeric value. For example, Sr^{2+} has a +2 charge and Br^{-} has a –1 charge.

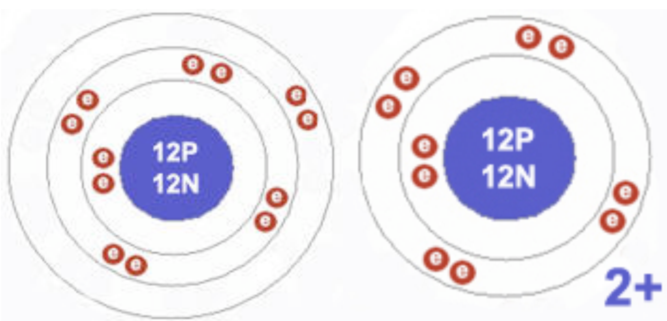


Figure 2.9 Diagram of Mg and Mg^{2+}

Cations and anions have physical and chemical properties that are very different than those of the elements from which they are formed. For example, elemental magnesium is a shiny metal that burns in air with a bright white flame. Magnesium ions are colorless and are found in most drinking water.

Most elements in the main groups of the periodic table (Groups 1A–7A) form monoatomic ions that have a charge related to the group number of the element.

- Metals in Groups 1A, 2A, and 3A form cations that have a positive charge equal to the group number of the element.
  Example: Sodium, Group 1A Na^{+} Calcium, Group 2A Ca^{2+}
- Nonmetals in Groups 5A, 6A, and 7A form anions that have a negative charge equal to 8 minus the group number of the element.
  Example: Oxygen, Group 8A O^{2-} Bromine, Group 7A Br^{-}

Other elements form ions with charges that are not easily predicted.

- Hydrogen forms both H^+ and H^- ions.
- Transition metals typically form cations with charges ranging from +1 to +3. Many transition metals form more than one monoatomic ion.
- Group 4A contains both metals and nonmetals, so some elements in this group form cations and others form anions.
- Other than aluminum, the metals in Groups 3A, 4A, and 5A form cations with positive charges that are not easily predicted.

Some of the more common monoatomic ions are shown in Figure 2.10. Notice that in general, metals form cations and nonmetals form anions and that the charges on most monoatomic ions are relatively small, between +2 and –2. Also, note that the noble gases (Group 8A) are quite unreactive and therefore do not form ions.

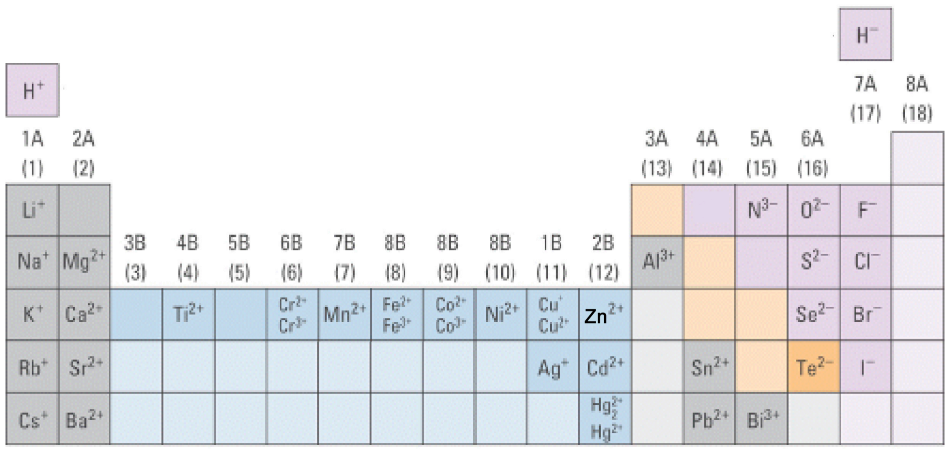


Figure 2.10 Charges on some monoatomic ions

EXAMPLE PROBLEM: Monoatomic Ions

(a) How many protons and electrons are in a Ca^{2+} ion?

(b) Identify the ion formed by phosphorus. How many protons and electrons are in this ion?

SOLUTION:

(a) Calcium is element 20 and the neutral atom has 20 protons and 20 electrons. A Ca^{2+} cation has lost two electrons and has 20 protons and 18 electrons.

(b) Phosphorus is in Group 5A and therefore forms an ion with a –3 charge. Phosphorus is element 15 and the P^{3-} ion has 15 protons and 18 electrons.

OWL Example Problems

2.15 Ions: Determine # Protons, Electrons: Tutor

2.16 Ions: Predicting Charge: Tutor

2.17 Monatomic Ions

Polyatomic Ions

Polyatomic ions are groups of covalently bonded atoms that carry an overall positive or negative charge. The formulas, names, and charges of the common polyatomic ions are shown in Table 2.8 and should be memorized. Most polyatomic ions are anions; there is only one common polyatomic cation, the ammonium ion (NH_4^+). You have seen many of the polyatomic ions in the formulas of inorganic acids. For example, HNO_3, nitric acid, contains the monoatomic ion H^+ and the polyatomic nitrate ion, NO_3^-.

Table 2.8 Names and formulas of common polyatomic ions

Cation (1+)			
NH_4^+	Ammonium		
Anions (1−)			
OH^-	Hydroxide	NO_2^-	Nitrite
HSO_4^-	Hydrogen sulfate	NO_3^-	Nitrate
CH_3COO^-	Acetate	MnO_4^-	Permanganate
ClO^-	Hypochlorite	$H_2PO_4^-$	Dihydrogen phosphate
ClO_2^-	Chlorite	CN^-	Cyanide
ClO_3^-	Chlorate	HCO_3^-	Hydrogen carbonate (bicarbonate)
ClO_4^-	Perchlorate		
Anions (2−)			
CO_3^{2-}	Carbonate	SO_3^{2-}	Sulfite
HPO_4^{2-}	Monohydrogen phosphate	SO_4^{2-}	Sulfate
$Cr_2O_7^{2-}$	Dichromate	$C_2O_4^{2-}$	Oxalate
$S_2O_3^{2-}$	Thiosulfate		
Anion (3−)			
PO_4^{3-}	Phosphate		

Chapter Goals Revisited

- Identify ionic compounds and their properties.
  Write formulas and names of ionic compounds.

OWL Example Problems
2.18 Names and Formulas of Polyatomic Ions: Tutor
2.19 Names and Formulas of Oxo Anions
2.20 Names and Formulas of Other polyatomic ions

Representing Ionic Compounds: Formulas

Ionic compounds are represented by empirical formulas that show the simplest ratio of cations and anions in the compound. In the formula of an ionic compound, the cation symbol or formula is always written first, followed by the anion symbol or formula. Ionic compounds do not have a positive or negative charge because the total cationic positive charge is balanced by the total anionic negative charge. This means that unlike covalent compounds, it is possible to predict the formula of an ionic compound if the cation and anion charges are known.

For example, the formula of the ionic formed from magnesium and bromine is predicted using the charges on the ions formed from these elements.

- Magnesium is in Group 2A and forms the Mg^{2+} ion.
- Bromine is in Group 7A and forms the Br^- ion.
- The +2 cation charge must be balanced by a –2 charge for the overall formula to carry no net charge. Two Br^- ions, each with a –1 charge, are needed to balance the Mg^{2+} ion charge. The formula of the ionic compound is therefore $MgBr_2$.

When the formula of an ionic compound contains a polyatomic ion, parentheses are used if more than one polyatomic ion is needed to balance the positive and negative charges in the compound. For example, the formula $Ca(NO_3)_2$ indicates that it contains Ca^{2+} ions and NO_3^- ions in a 1:2 ratio.

EXAMPLE PROBLEM: Ionic Compound Formulas

(a) What is the formula of the ionic compound expected to form between the elements oxygen and sodium?

(b) What is the formula of the ionic compound formed between the ions Zn^{2+} and PO_4^{3-}?

(b) What ions make up the ionic compound $Cr(NO_3)_3$?

SOLUTION:

(a)	Na_2O	Sodium is in Group 1A and forms a cation with a +1 charge, Na^+. Oxygen is in group 6A and forms an anion with a –2 charge, O^{2-}. Two Na^+ ions are required to provide a total +2 positive charge that balances the –2 charge on O^{2-}.
(b)	$Zn_3(PO_4)_2$	In this case, more than one of each ion is needed in order to balance the positive and negative charges. Three Zn^{2+} ions provide a positive charge of +6, and two PO_4^{3-} ions provide a negative charge of –6. Parentheses are used to indicate the total number of polyatomic ions in the compound formula.
(c)	Cr^{3+}, NO_3^-	Cr is a transition metal and it is not possible to predict its charge when it forms an ion. NO_3^- is a polyatomic ion with a –1 charge. The three NO_3^- ions in the compound formula provide a total negative charge of –3, so the single cation must have a +3 charge to balance this negative charge.

OWL Example Problems
2.21 Ionic Compound Formulas

Naming Ionic Compounds

Like covalent compounds, ions and ionic compounds are named using guidelines created by IUPAC, the International Union of Pure and Applied Chemistry. Naming ionic compounds involves identifying the charges on the monoatomic and polyatomic ions in a chemical formula so it is very important to memorize the rules for predicting the charges on monoatomic ions and the names, formulas, and charges on polyatomic ions.

Ions and ionic compounds are named according to the rules shown below.

Monoatomic cations

- The name of a main group monoatomic cation is the element name followed by the word *ion*.
  Example: Na^+ sodium ion Mg^{2+} magnesium ion
- The name of a transition metal cation is the element name followed by the cation charge in Roman numerals within parentheses and the word *ion*.
  Example: Fe^{2+} iron(II) ion Co^{3+} cobalt(III) ion

Monoatomic anions

- The name of a monoatomic anion is the element name changed to include the suffix *–ide*, followed by the word *ion.*
  Example: Br^- bromide ion O^{2-} oxide ion

Polyatomic ions

- The name of polyatomic cations and anions are shown in Table 2.8 and must be memorized. Notice that the names also include the word *ion.*
  Example: NO_3^- nitrate ion MnO_4^- permanganate ion

Ionic Compounds

- The name of an ionic compounds consists of the cation name followed by the anion name. The word *ion* is dropped because the compound does not carry a charge. Prefixes are not used to indicate the number of ions present in the formula of an ionic compound.
  Example: $NaNO_3$ sodium nitrate Co_2O_3 cobalt(III) oxide

EXAMPLE PROBLEM: Ionic Compound Names

(a) What is the name of the compound with the formula CuCN?
(b) What is the formula for aluminum nitrite?

SOLUTION:

(a) Copper(I) cyanide — The cation is a transition metal and its name must include the cation charge. The cyanide ion, CN^-, is a polyatomic ion with a –1 charge. The single copper cation therefore has a +1 charge. The name of this compound includes the charge on the cation in Roman numerals, within parentheses.

(b) $Al(NO_2)_3$ — Aluminum is in Group 3A and forms a cation with a +3 charge, Al^{3+}. The nitrite ion, NO_2^-, is a polyatomic ion whose name, charge, and formula must be memorized. Three nitrite ions are needed to provide a total –3 charge that balances the +3 charge on Al^{3+}. Parentheses are used to indicate the total number of polyatomic ions in the compound formula.

OWL Example Problems
2.23 Naming Ionic Compounds: Tutor
2.22 Naming Ionic Compound

Identifying Covalent and Ionic Compounds

It is often challenging to determine if a compound is covalent and ionic. However, the formula or name of a compound contains information that can be used to classify a compound. The following guidelines show how the two classes of compounds are similar and how they differ.

Covalent compounds

- Contain only nonmetals
- Are named using prefixes to indicate the number of each element in a formula
  Examples: CO_2, carbon dioxide; H_2O, water; N_2O_5, dinitrogen pentaoxide

Ionic compounds

- Contain monoatomic and/or polyatomic ions
- Usually contain metals and nonmetals but can also contain only nonmetals
- Are never named using prefixes and sometimes with the cation charge in Roman numerals within parentheses
  Examples: NaF, sodium fluoride; $(NH_4)_2CO_3$, ammonium carbonate;
  CuO, copper(II) oxide

Chapter Review

OWL Summary Assignments
2.x Chapter Review
2.x Challenge Problems

Key Equations

$$\text{Average atomic mass} = \sum_{\substack{\text{all}\\ \text{isotopes}}} (\text{exact mass})(\text{fractional abundance}) \qquad \textbf{(2.1)}$$

Key Terms

2.1 The Structure of the Atom
proton
neutron
electron
atomic nucleus
atomic mass unit (u)
ion
cation
anion
atomic number (Z)
mass number (A)
atomic symbol
isotopes
atomic mass

2.2 Elements and the Periodic Table
periodic table of the elements
groups
periods
main group
transition elements
lanthanides
actinides
metals
nonmetals
metalloids
semi-metals
allotropes

2.3 Covalent Compounds
covalent compounds
molecular covalent compound
network covalent compound
molecular formula
empirical formula
structural formula
condensed structural formula
wedge and dash model
ball and stick model
space filling model
binary nonmetals
inorganic acids
hydrocarbons
oxoacids

2.4 Ions and Ionic Compounds
ionic compounds
monoatomic ion
polyatomic ions